CHAPTER 17 Properties of Solutions

Note: Before working the problems of this chapter, you should familiarize yourself with table 17.1, p. 316, Table 17.2, p. 330, Tables 17.3 and 17.4, p. 332, and Table 17.5, p. 333.

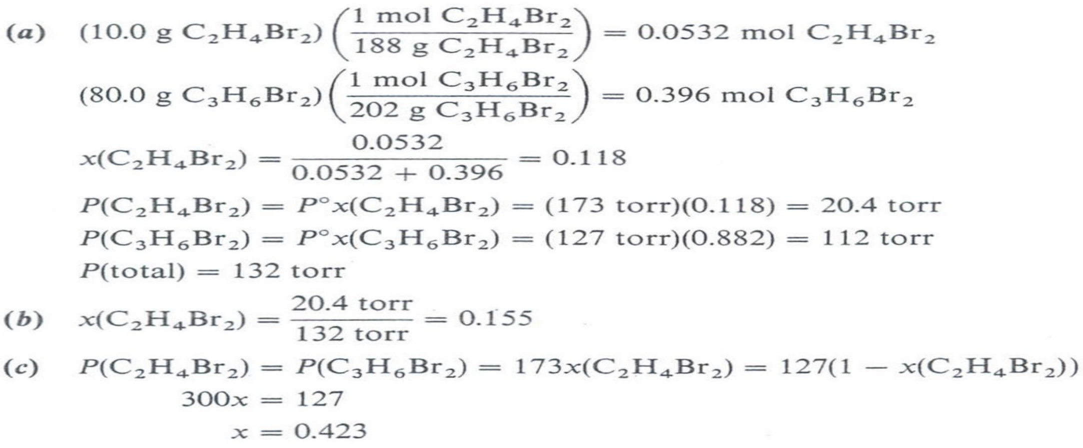
17.1 RAOULT'S LAW AND VAPOR PRESSURE LOWERING

17.1 At 40°C the vapor pressure, in torr, of methyl alcohol-ethyl solutions is represented by

P = 119x + 135 where x is the mole fraction of methyl alcohol. What are the vapor pressures of the pure components at this temperature?

- If x = 0, P = 135 torr
  If x=1, P = 119 + 135 = 254 torr

17.2 Ethylene bromide, $C_2H_4Br_2$, and 1,2-dibromopropane, $C_3H_6Br_2$, form a series of ideal solutions over the whole range of composition. At 85°C the vapor pressures of these two pure liquids are 173 and 172 toee, respectively. (a) If 10.0 g of ethylene bromide is dissolved in 80.0 g of 1,2-dibromopropane, calculate the partial pressure of each component and the total pressure of the solution at 85°C. (b) Calculate the mole fraction of ethylene bromide in thee vapor in equilibrium with the above solution. (c) What would be the mole fraction of ethylene bromide in a solution at 85°C equilibrated with a 50:50 mole mixture in the vapor?

- 

(a)
$$(10.0\text{ g } C_2H_4Br_2)\left(\frac{1\text{ mol } C_2H_4Br_2}{188\text{ g } C_2H_4Br_2}\right) = 0.0532\text{ mol } C_2H_4Br_2$$
$$(80.0\text{ g } C_3H_6Br_2)\left(\frac{1\text{ mol } C_3H_6Br_2}{202\text{ g } C_3H_6Br_2}\right) = 0.396\text{ mol } C_3H_6Br_2$$
$$x(C_2H_4Br_2) = \frac{0.0532}{0.0532 + 0.396} = 0.118$$
$$P(C_2H_4Br_2) = P^\circ x(C_2H_4Br_2) = (173\text{ torr})(0.118) = 20.4\text{ torr}$$
$$P(C_3H_6Br_2) = P^\circ x(C_3H_6Br_2) = (127\text{ torr})(0.882) = 112\text{ torr}$$
$$P(\text{total}) = 132\text{ torr}$$

(b)
$$x(C_2H_4Br_2) = \frac{20.4\text{ torr}}{132\text{ torr}} = 0.155$$

(c)
$$P(C_2H_4Br_2) = P(C_3H_6Br_2) = 173x(C_2H_4Br_2) = 127(1 - x(C_2H_4Br_2))$$
$$300x = 127$$
$$x = 0.423$$

17.3 The vapor pressure of pure liquid solvent A is 0.80 atm. When a nonvolatile substance B is added to the solvent, its vapor pressure drops to 0.60 atm. What is the mole fraction of component B in the solution?

- $P = xP^\circ \qquad x(\mathrm{A}) = \frac{P}{P^\circ} = \frac{0.60\ \mathrm{atm}}{0.80\ \mathrm{atm}} = 0.75 \qquad x(\mathrm{B}) = 1 - x(\mathrm{A}) = 1 - 0.75 = 0.25$

17.4 The vapor pressure of pure water at 26 °C is 25.21 torr. What is the vapor pressure of a solution which contains 20.0 g glucose, $C_6H_{12}O_6$, in 70 g water?

$$\frac{20.0\ \mathrm{g\ glucose}}{180\ \mathrm{g/mol}} = 0.111\ \mathrm{mol\ glucose} \qquad \frac{70.0\ \mathrm{g\ H_2O}}{18.0\ \mathrm{g/mol}} = 3.89\ \mathrm{mol\ H_2O}$$

$$x(\mathrm{H_2O}) = \frac{3.89}{0.111 + 3.89} = 0.972$$

$$P = P^\circ x(\mathrm{H_2O}) = (25.21\ \mathrm{torr})(0.972) = 24.5\ \mathrm{torr}$$

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17.5 The vapor pressure of pure water at 25 °C is 23.76 torr. The vapor pressure of a solution containing 5.40 g of a nonvolatile substance in 90.0 g water is 23.32 torr. Compute the molecular weight of the solute.

$$P = P^\circ x$$

- $x = \frac{23.32\ \mathrm{torr}}{23.76\ \mathrm{torr}} = 0.981 = \frac{5.00}{5.00 + z}$

$$(5.00 + z)(0.981) = 5.00$$

$$z = 0.0968\ \mathrm{mol}$$

$$\frac{5.40\ \mathrm{g}}{0.0968\ \mathrm{mol}} = 55.8\ \mathrm{g/mol}$$

17.6 At 20 °C the vapor pressure of methyl alcohol (CH_3OH) is 94 torr and the vapor pressure of ethyl alcohol (C_2H_5OH) is 44 torr. Being closely related, these compounds form a two-component system which adheres quite closely to Raoult’s law throughout the entire range of concentrations. If 20 g of CH_3OH is mixed with 100 g of C_2H_5OH, determine the partial pressure exerted by each and the total pressure of the solution. Calculate the composition of the vapor above the solution by applying Dalton’s law.

- In an ideal solution of two liquids, there is no distinction between solute and solvent, and Raoult’s law holds for each component of such solutions. Hence, when two liquids are mixed to give an ideal solution, the partial pressure of each liquid is equal to its vapor pressure multiplied by its mole fraction in the solution. The molecular weights of CH_3OH and C_2H_5OH are 32 and 46, so

$$\text{Partial pressure of } CH_3OH = (94\text{ torr})\left[\frac{\frac{20}{32}\text{ mol } CH_3OH}{\frac{20}{32}\text{ mol } CH_3OH + \frac{100}{46}\text{ mol } C_2H_5OH}\right] = (94\text{ torr})(0.22) = 21\text{ torr}$$

$$\text{Partial pressure of } C_2H_5OH = (44\text{ torr})\left[\frac{\frac{100}{46}\text{ mol } C_2H_5OH}{\frac{20}{32}\text{ mol } CH_3OH + \frac{100}{46}\text{ mol } C_2H_5OH}\right] = (44\text{ torr})(0.78) = 34\text{ torr}$$

The total pressure of the gaseous mixture is the sum of the partial pressures of all the components (Dalton’s law): Total pressure of solution = (21+34) torr = 55 torr. Dalton’ s law also incidates that the mole fraction of any component of a gaseous mixture is equal to its pressure fraction, i.e., its partial pressure divided by the total pressure.

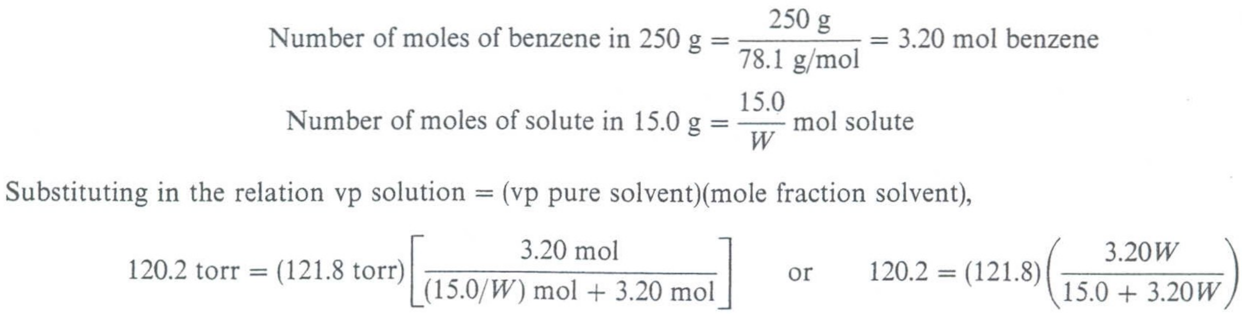
$$\text{Mole fraction of } CH_3OH \text{ in vapor} = \frac{\text{partial pressure of } CH_3OH}{\text{total pressure}} = \frac{21\text{ torr}}{55\text{ torr}} = 0.38$$

$$\text{Mole fraction of } C_2H_5OH \text{ in vapor} = \frac{\text{partial pressure of } C_2H_5OH}{\text{total pressure}} = \frac{34\text{ torr}}{55\text{ torr}} = 0.62$$

Since the mole fraction for (ideal) gases is the same as the volume fraction, we may also say that the vapor consists 0f 38% CH_3OH by volume. Note that the vapor is relatively richer in the more volatile component, methyl alcohol (mole fraction, 0.38), than is the liquid (mole fraction of CH_3OH, 0.22).

17.7 At 30 °C, pure benzene (molecular weight 78.1 g/mol) has a vapor pressure of 121.8 torr. Dissolving 15.0 g of a nonvolatile solute in 250 g of benzene produced a solution having a vapor pressure of 120.2 torr. Determine the approximate molecular weight of the solute.

- Let W be the molecular weight of the solute.



$$\text{Number of moles of benzene in 250 g} = \frac{250\text{ g}}{78.1\text{ g/mol}} = 3.20\text{ mol benzene}$$

$$\text{Number of moles of solute in 15.0 g} = \frac{15.0}{W}\text{ mol solute}$$

Substituting in the relation vp solution = (vp pure solvent)(mole fraction solvent),

$$120.2\text{ torr} = (121.8\text{ torr})\left[\frac{3.20\text{ mol}}{(15.0/W)\text{ mol} + 3.20\text{ mol}}\right] \qquad \text{or} \qquad 120.2 = (121.8)\left(\frac{3.20W}{15.0 + 3.20W}\right)$$

17.78 The vapor pressure of water at 28°C is 28.35 torr. Compute the vapor pressure at 28°C of a solution containing 68 g of sugar cane, $C_{12}H_{22}O_{11}$, in 1000 g of water.

$$\text{Mol of } C_{12}H_{22}O_{11} \text{ in 68 g} = \frac{68\text{ g}}{342\text{ g/mol}} = 0.20 \text{ mol } C_{12}H_{22}O_{11}$$

$$\text{Mol of } H_2O \text{ in 1000 g} = \frac{1000\text{ g}}{18.02\text{ g/mol}} = 55.49 \text{ mol } H_2O$$

$$\text{Total mol} = (0.20 + 55.49)\text{ mol} = 55.69\text{ mol}$$

$$\text{Mole fraction } C_{12}H_{22}O_{11} = \frac{0.20}{55.69} = 0.0036 \qquad \text{Mole fraction } H_2O = \frac{55.49}{55.69} = 0.9964$$

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First method

Vapor pressure of solution = (vp of pure solvent)(mole fraction of solvent) = (28.35 torr)(0.9964) = 28.25 torr

Second method

Vapor pressure depression = ΔP = (vp of pure solvent)(mole fraction of solute) = (28.35 torr)(0.0036) = 0.10 torr

Vapor pressure of solution = (28.35 − 0.10) torr = 28.25 torr

17.9 Calculate the mole fraction of toluene in the vapor phase which is in equilibrium with a solution of benzene and toluene having a mole fraction of toluene 0.500. the vapor pressure of pure benzene is 119 torr; that of toluene is 37.0 torr at the same temperature.

Assuming ideal behavior,

$$P_b = x_b P_b^\circ = 0.500(119\ \text{torr}) = 59.5\ \text{torr}$$
$$P_t = x_t P_t^\circ = 0.500(37.0\ \text{torr}) = 18.5\ \text{torr}$$
$$P_{\text{tot}} = 78.0\ \text{torr}$$
$$\text{Mole fraction toluene} = \frac{18.5\ \text{torr}}{78.0\ \text{torr}} = 0.237$$

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17.10 What is the composition of the vapor which is in equilibrium at 30 °C with a benzene-toluene solution with a mole fraction of benzene of 0.400? with a mole fraction of benzene of 0.6000? $P_b^o = 119\ torr\ \ and\ P_t^o = 37.0\ torr$

$$P_b = (0.400)(119\ \text{torr}) = 47.6\ \text{torr} \qquad P_t = (0.600)(37.0\ \text{torr}) = 22.2\ \text{torr} \qquad P_{\text{tot}} = (47.6 + 22.2)\ \text{torr} = 69.8\ \text{torr}$$

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The composition of the vapor is determined by applying Dalton's law of partial pressures:

$$x_b = \frac{P_b}{P_{\text{tot}}} = \frac{47.6\ \text{torr}}{69.8\ \text{torr}} = 0.682 \qquad x_t = \frac{22.2\ \text{torr}}{69.8\ \text{torr}} = 0.318 = 1.000 - 0.682$$

Similarly, for the case of the solution in which the mole fraction of toluene is 0.400,

$$x_b = \frac{71.5}{86.3} = 0.829 \qquad x_t = \frac{14.8}{86.3} = 0.171$$

17.11 At 30°C, the vapor pressure of pure benzene, C_6H_6, is 119 torr, while that of toluene, $C_6H_5CH_3$, is 37 torr. Assuming ideal behavior, plot a vapor pressure-composition diagram for solutions of the two at 30°C. Include on the diagram the compositions of vapor in equilibrium with the solutions.

- See Fig. 17.1. The vapor pressure of each component is calculated at several different points, and a smooth curve is drawn through them to represent the pressure-composition curve of the vapor. (The corresponding curvefor the liquid phase in an ideal solution is a straight line connecting the vapor pressure of the pure components.) An example is given:

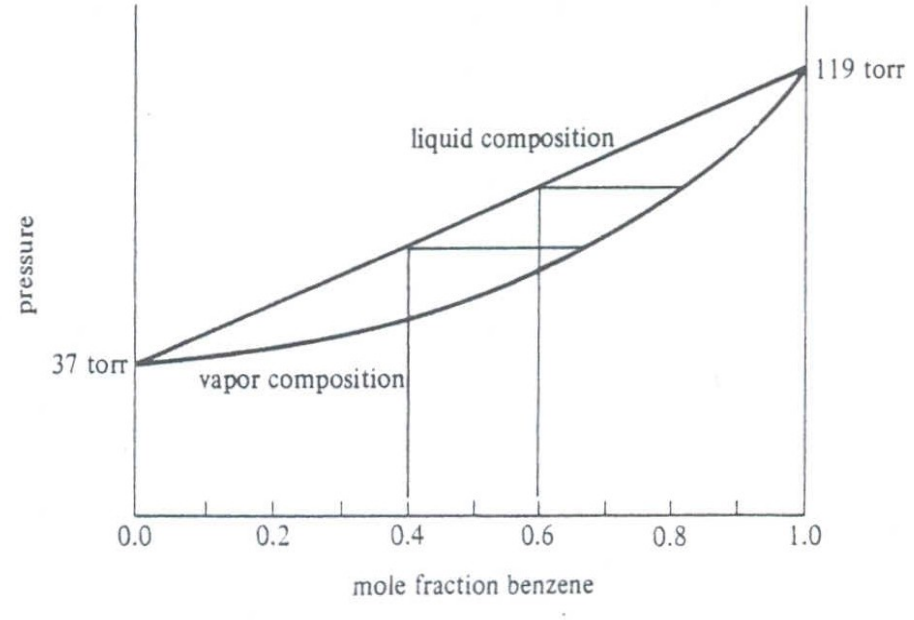


Fig. 17.1

At

$$x_b = 0.25 \qquad P_b = (0.25)(119\text{ torr}) = 30\text{ torr}$$
$$x_t = 0.75 \qquad P_t = (0.75)(37\text{ torr}) = 28\text{ torr}$$
$$P_{tot} = 58\text{ torr}$$

In the vapor phase,

$$x_b = \frac{30\text{ torr}}{58\text{ torr}} = 0.52$$

17.12 At 50 °C the vapor pressure of pure CS_2 is 854 torr. A solution of 2.0 g of sulfur in 100 g of CS_2 has a vapor pressure 848.9 torr. Determine the formula of the sulfur molecule.

- The mole fraction of CS_2 is determined from Raoult's law:

$$x = \frac{P}{P°} = \frac{848.9\text{ torr}}{854\text{ torr}} = 0.994$$

The mole fraction of the sulfur molecules, S_i, is thus 0.006. The molality of S_i is

$$\frac{0.006\text{ mol}}{(0.994\text{ mol }CS_2)(76\text{ g/mol})(1\text{ kg}/10^3\text{ g})} = 0.079\text{ m}$$

Hence 0.079 mol S_i/kg CS_2 is 20 g S_i/kg CS_2; hence the molecular weight of S_i is

$$\frac{20\text{ g}}{0.079\text{ mol}} = 250\text{ g/mol}, \qquad \text{so} \qquad x = \frac{250}{32} = 8 \qquad \text{The formula is } S_8.$$

17.13 Calculate the vapor pressure lowering of a 0.100 m aqueous solution of nonelectrolyte at 75 °C.

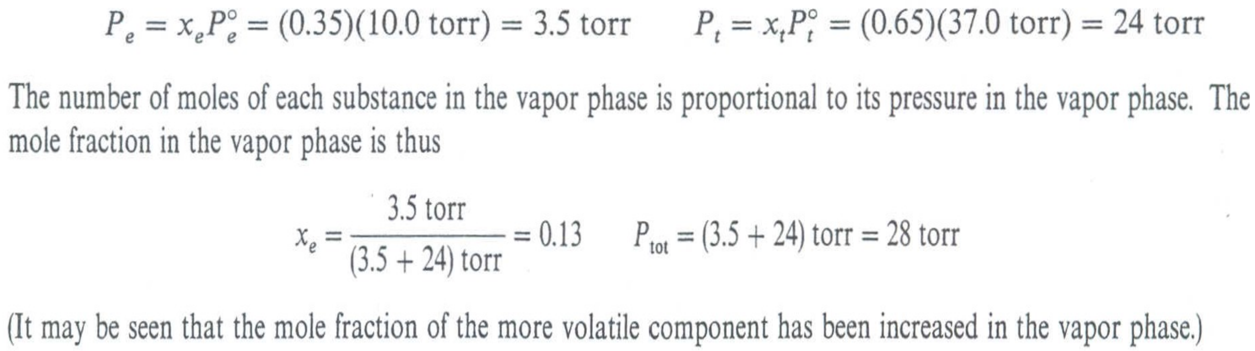
- The vapor pressure of pure water, P_A°, at 75 °C is determined as follows:

$$\log\frac{P_2}{P_1} = \frac{\Delta H°}{2.30}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) = \frac{9720\text{ cal}}{2.30(1.987\text{ cal/mol}\cdot\text{K})}\left(\frac{1}{348} - \frac{1}{373}\right) = 0.410$$
$$P_2/P_1 = 2.57 \qquad P_1 = 296\text{ torr}$$

The vapor pressure lowering due to 0.100 m solute is

$$\Delta P = mM_A P_A^\circ = \left(\frac{0.100\text{ mol}}{1000\text{ g}}\right)\left(\frac{18.0\text{ g}}{\text{mol}}\right)(296\text{ torr}) = 0.533\text{ torr}$$

17.14 Calculate the composition of the vapor in equilibrium with an ideal solution of ethylbenzene ($P_e^o = 10.0\ torr$) and methylbenzene ($P_m^o = 37.0\ torr$) in



$$P_e = x_e P_e^\circ = (0.35)(10.0\text{ torr}) = 3.5\text{ torr} \qquad P_t = x_t P_t^\circ = (0.65)(37.0\text{ torr}) = 24\text{ torr}$$

The number of moles of each substance in the vapor phase is proportional to its pressure in the vapor phase. The mole fraction in the vapor phase is thus

$$x_e = \frac{3.5\text{ torr}}{(3.5 + 24)\text{ torr}} = 0.13 \qquad P_{\text{tot}} = (3.5 + 24)\text{ torr} = 28\text{ torr}$$

(It may be seen that the mole fraction of the more volatile component has been increased in the vapor phase.)

which the mole fraction of ethylbenzene in the liquid is 0.35. Calculate the total vapor pressure of the solution.

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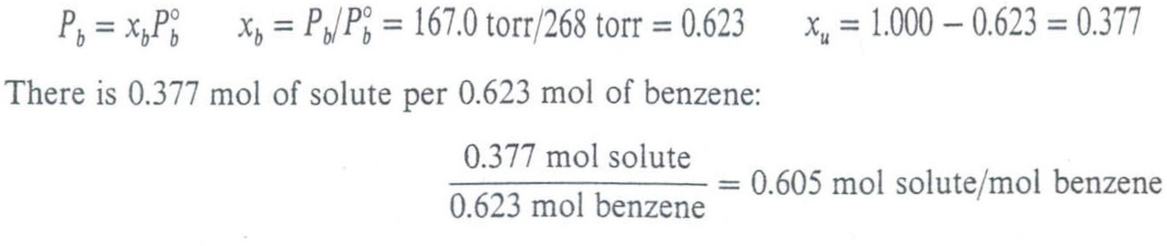
17.15 Thee vapor pressure of pure benzene, C_6H_6, at 50 °C is 268 torrr. How many mol of nonvolatile solute per mol of benzene is required to prepare a solution of benzene having a vapor pressure of 167.0 torr at 50 °C?

- 

$$P_b = x_b P_b^\circ \qquad x_b = P_b/P_b^\circ = 167.0\text{ torr}/268\text{ torr} = 0.623 \qquad x_u = 1.000 - 0.623 = 0.377$$

There is 0.377 mol of solute per 0.623 mol of benzene:

$$\frac{0.377\text{ mol solute}}{0.623\text{ mol benzene}} = 0.605\text{ mol solute/mol benzene}$$

17.16 At 25 °C, the vapor pressure of methyl alcohol, CH_3OH, is 96.0 torr. What is the mole fraction of CH_3OH in a solution in which the (partial) vapor pressure of CH_3OH is 23.0 torr at 25 °C?

- $$x = \frac{P}{P^\circ} = \frac{23.0\ \text{torr}}{96.0\ \text{torr}} = 0.240$$

17.17 At 25 °C, the vapor pressure of pure benzene is 100 torr, while that of pure ethyl alcohol is 44 torr. Assuming ideal behavior, calculate the vapor pressure at 25 °C of a solution which contains 10.0 g of each substance.

- $$(10.0\ \text{g}\ C_6H_6)\left(\frac{1\ \text{mol}}{78.0\ \text{g}}\right) = 0.128\ \text{mol}\ C_6H_6 \qquad (10.0\ \text{g}\ C_2H_5OH)\left(\frac{1\ \text{mol}}{46.0\ \text{g}}\right) = 0.217\ \text{mol}\ C_2H_5OH$$
  $$x_b = \frac{0.128\ \text{mol}}{(0.128 + 0.217)\ \text{mol}} = 0.371 \qquad x_a = 1.000 - 0.371 = 0.629$$
  $$P_b = x_b P_b^\circ = (0.371)(100\ \text{torr}) = 37.1\ \text{torr} \qquad P_a = x_a P_a^\circ = (0.629)(44\ \text{torr}) = 28\ \text{torr}$$
  $$P_{\text{tot}} = 65\ \text{torr}$$

17.18 Estimate the lowering of the vapor pressure due to the solute in a 1.0 m aqueous at 100 °C.

The normal boiling point of water is 100 °C; hence

$$P^\circ(H_2O) = 1.00\ \text{atm} \qquad M(H_2O) = 18.0\ \text{g/mol}$$
$$\Delta P = mM_A P_A^\circ = \left(\frac{1.0\ \text{mol}}{1000\ \text{g}}\right)\left(\frac{18.0\ \text{g}}{\text{mol}}\right)(1.0\ \text{atm}) = 0.018\ \text{atm} = 14\ \text{torr}$$

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17.2 FREEZING POINT DEPRESSION AND BOILING POINT ELEVATION

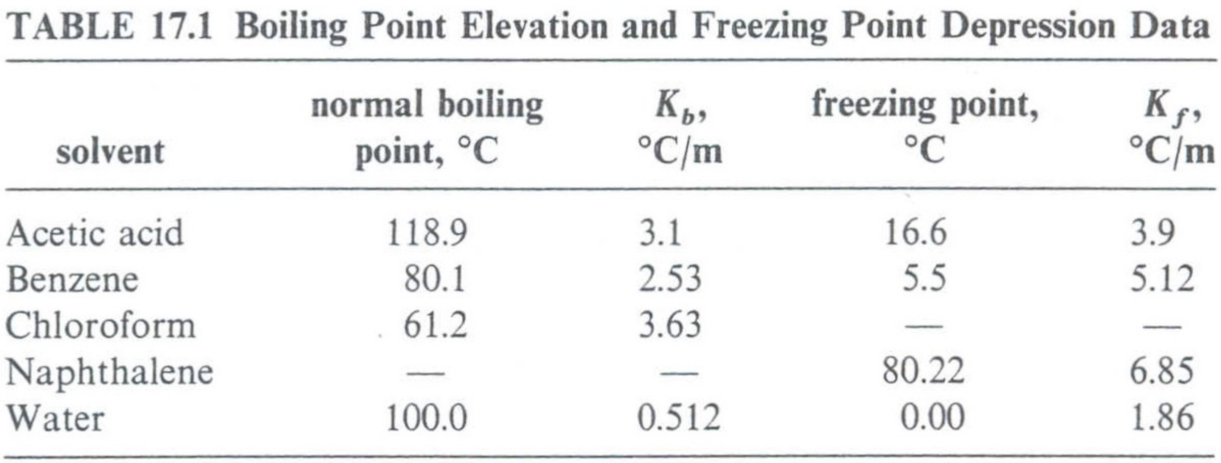


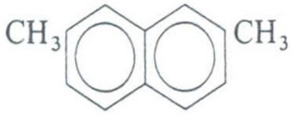
TABLE 17.1 Boiling Point Elevation and Freezing Point Depression Data

solvent	normal boiling point, °C	K_b, °C/m	freezing point, °C	K_f, °C/m
Acetic acid	118.9	3.1	16.6	3.9
Benzene	80.1	2.53	5.5	5.12
Chloroform	61.2	3.63	—	—
Naphthalene	—	—	80.22	6.85
Water	100.0	0.512	0.00	1.86

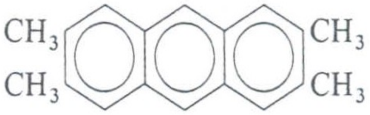
In the problems that follow, temperature on the Celsius scale will be denoted t, with T denoting (as previously) absolute temperature. Recall that the two scales utilize the same size of a degree so that for differences in temperature: 1 °C = 1K.

17.19 Calculate the freezing point of the solution containing 15.6 g of solute/kg of benzene for each of the following solutes:

(*a*)



(*b*)



- (*a*) $(15.6\ \text{g})\left(\frac{1\ \text{mol}}{156\ \text{g}}\right) = 0.100\ \text{mol}$

Since the solute is dissolved in 1 kg of solvent, the molality is 0.100 m.

$$\Delta t = (0.100\ \text{m})(4.9\ ^\circ\text{C/m}) = 0.49\ ^\circ\text{C} = 5.5^\circ - t$$
$$t = (5.5 - 0.49)\ ^\circ\text{C} = 5.0\ ^\circ\text{C}$$

(*b*) $(15.6\ \text{g})\left(\frac{1\ \text{mol}}{234\ \text{g}}\right) = 0.0667\ \text{mol}$

In 1.00 kg of benzene:

$$\Delta t = (0.0667\ \text{m})(4.9\ ^\circ\text{C/m}) = 0.33\ ^\circ\text{C} = 5.5^\circ - t$$
$$t = (5.5 - 0.33)\ ^\circ\text{C} = 5.2\ ^\circ\text{C}$$

17.20 What is the freezing point of a 10% (by weight) solution of CH_3OH in water?

1.00 kg contains 0.100 kg CH_3OH (and 0.900 kg H_2O).

$$(100\text{ g }CH_3OH)\left(\frac{1\text{ mol}}{32.0\text{ g}}\right) = 3.12\text{ mol} \qquad \frac{3.12\text{ mol}}{0.900\text{ kg}} = 3.47\text{ m}$$

$$\Delta t = Km = (1.86\ °\text{C/m})(3.47\text{ m}) = 6.45\ °\text{C}$$

$$\text{fp} = -6.45\ °\text{C}$$

17.21 When 10.6 g of a nonvolatile substance is dissolved in 740 g of ether, its boiling point is raised 0.284 °C. What is the molecular weight of the substance? Modal boiling-point constant for ether is 2.11 $°\text{C} \cdot kg$/mol.

$$\Delta t = Km = 0.284\ °\text{C}$$

$$m = \frac{0.284\ °\text{C}}{2.11\ °\text{C/m}} = 0.135\text{ m} = \frac{0.135\text{ mol}}{\text{kg}}$$

$$\frac{10.6\text{ g}}{0.740\text{ kg ether}} = \frac{14.3\text{ g}}{\text{kg ether}} \qquad \frac{14.3\text{ g/kg}}{0.135\text{ mol/kg}} = 106\text{ g/mol}$$

17.22 The freezing point of a sample of naphthalene was found to be 80.6 °C. When 0.512 g of a substance is dissolved in 7.03 g naphthalene, the solution has freezing point of 75.2 °C. What is the molecular weight of the solute? The molal freezing point constant of naphthalene is 6.80 $°\text{C} \cdot kg$/mol.

$$\Delta t_f = (80.6 - 75.2)\ °\text{C} = 5.4\ °\text{C} = K_f m$$

$$m = \frac{5.4\ °\text{C}}{6.80\ °\text{C/m}} = 0.79\text{ m} = \frac{0.79\text{ mol}}{\text{kg solvent}}$$

$$\frac{0.512\text{ g}}{0.00703\text{ kg}} = \frac{72.8\text{ g}}{\text{kg}} \qquad \frac{72.8\text{ g/kg}}{0.79\text{ mol/kg}} = 92\text{ g/mol}$$

17.23 If glycerin, $C_3H_5(OH)_3$, and methyl alcohol, CH_3OH, sell at the same price per pound, which would be cheaper for preparing an antifreeze solution for the radiator of an automobile?

- A given mass of methyl alcohol, because it has a lower molecular weight, contains more moles than the same mass of glycerin. More moles of solute means higher molality and thus lower freezing point. It would take less methyl alcohol to protect a radiator to a given freezing point. (Practically, however, methyl alcohol vaporizes too readily to be used in modern, high-temperature auto radiators.)

17.24 How much ethyl alcohol, C_2H_5OH, must be added to 1.00 L of water so that the solution will not freeze at -4^{0}F?

- $$-4\ ^\circ\text{F} = \tfrac{5}{9}(-4-32)\ ^\circ\text{C} = -20\ ^\circ\text{C}$$
  $$\Delta t = 20\ ^\circ\text{C} = Km = (1.86\ ^\circ\text{C/m})(m)$$
  $$m = \frac{20\ ^\circ\text{C}}{1.86\ ^\circ\text{C/m}} = 10.7\ \text{m} \qquad (10.7\ \text{mol})\left(\frac{46.0\ \text{g}}{\text{mol}}\right) = 495\ \text{g}$$

17.25 If the radiator of an automobile contains 12 L of water, how much would the freezing point be lowered by the addition of 5 kg of Prestone (glycol, $C_2H_4(OH)_2$)? How many kg of Zerone (methyl alcohol, CH_3OH) would be required to produce the same result? Assume 100 % purity.

$$(5.0\ \text{kg}\ C_2H_4(OH)_2)\left(\frac{10^3\ \text{g}}{\text{kg}}\right)\left(\frac{1\ \text{mol}}{62.0\ \text{g}}\right) = 80.6\ \text{mol} \qquad \frac{80.6\ \text{mol}}{12\ \text{kg}} = 6.7\ \text{m}$$
$$\Delta t = (1.86\ ^\circ\text{C/m})(6.7\ \text{m}) = 12\ ^\circ\text{C} \qquad \text{fp} = -12\ ^\circ\text{C}$$

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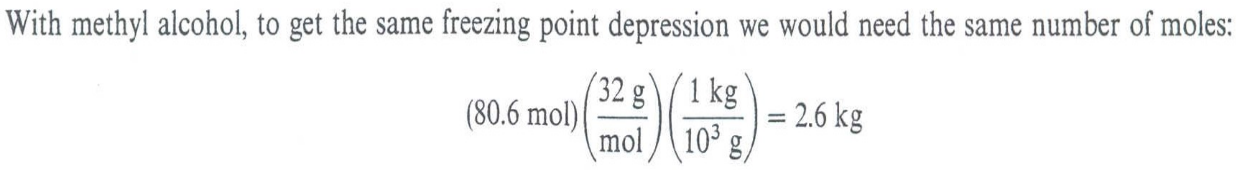
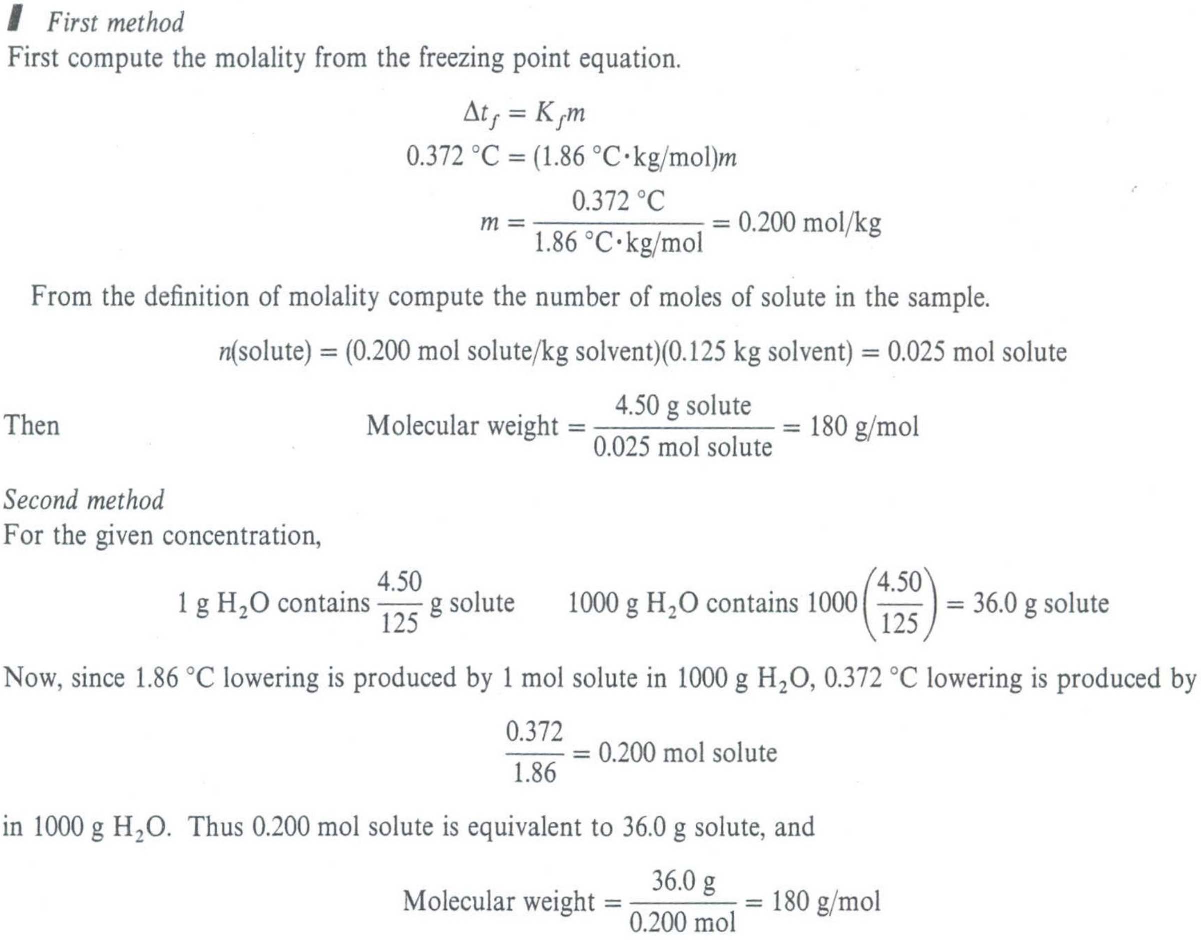


With methyl alcohol, to get the same freezing point depression we would need the same number of moles:

$$(80.6\ \text{mol})\left(\frac{32\ \text{g}}{\text{mol}}\right)\left(\frac{1\ \text{kg}}{10^3\ \text{g}}\right) = 2.6\ \text{kg}$$

17.26 What is the freezing point of a 10% (by weight) solution of CH_3OH in water?

- 

  First method
  First compute the molality from the freezing point equation.

  $$\Delta t_f = K_f m$$
  $$0.372\ ^\circ\text{C} = (1.86\ ^\circ\text{C}\cdot\text{kg/mol})m$$
  $$m = \frac{0.372\ ^\circ\text{C}}{1.86\ ^\circ\text{C}\cdot\text{kg/mol}} = 0.200\ \text{mol/kg}$$

  From the definition of molality compute the number of moles of solute in the sample.

  $$n(\text{solute}) = (0.200\ \text{mol solute/kg solvent})(0.125\ \text{kg solvent}) = 0.025\ \text{mol solute}$$

  Then $$\text{Molecular weight} = \frac{4.50\ \text{g solute}}{0.025\ \text{mol solute}} = 180\ \text{g/mol}$$

  Second method
  For the given concentration,

  $$1\ \text{g}\ H_2O\ \text{contains}\ \frac{4.50}{125}\ \text{g solute} \qquad 1000\ \text{g}\ H_2O\ \text{contains}\ 1000\left(\frac{4.50}{125}\right) = 36.0\ \text{g solute}$$

  Now, since 1.86 °C lowering is produced by 1 mol solute in 1000 g H_2O, 0.372 °C lowering is produced by

  $$\frac{0.372}{1.86} = 0.200\ \text{mol solute}$$

  in 1000 g H_2O. Thus 0.200 mol solute is equivalent to 36.0 g solute, and

  $$\text{Molecular weight} = \frac{36.0\ \text{g}}{0.200\ \text{mol}} = 180\ \text{g/mol}$$